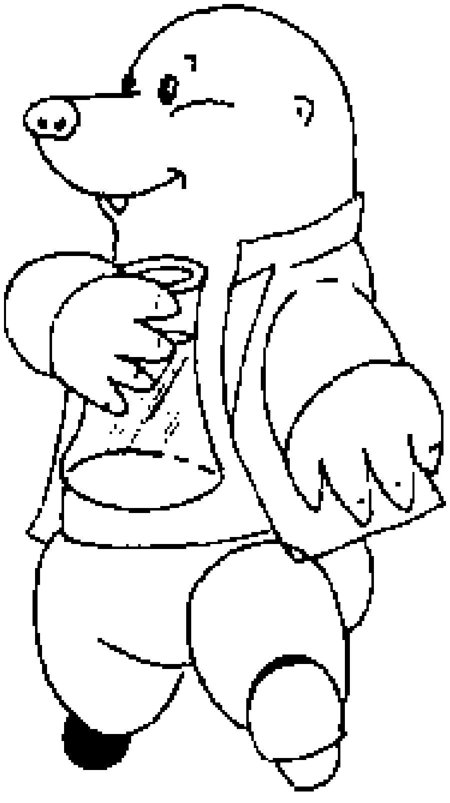
Chemistry I

Chemical Bonding



Why chemical bonding occurs

- Chemical bonding occurs:
  - So atoms can achieve maximum stability
    - By reducing potential energy
    - By reaching Noble gas configuration
- All bonding forces are due to electrostatic charge.
  - Opposite charges attract, Like charges repel.

Metallic bonding

- occurs between atoms of the same metal
- Atoms have 1,2 or 3 valence electrons, therefore there are many *vacancies in valence shell.*
- Nuclei of the atoms are in a sea of electrons
- When electron clouds overlap, *electrons can move into electron cloud of adjoining atoms*

Properties of substances with metallic bonding

- 1) Metals are dense because the atoms in metals are tightly packed in the lattice.

- 2) Metals have high melting and boiling points because strong forces of attraction exist between particles and these forces operate throughout the crystal.

Properties of substances with metallic bonding

- 3) Metals are good conductors of heat because delocalised electrons transmit the energy to its neighbors.
- 4) Metals are good conductors of electricity because electrons can flow within the lattice or crystal

Properties of substances with metallic bonding

- 5) Metals are malleable and ductile because bending and stretching does not disrupt the metallic bonds.
- 6) Metals are lustrous(shiny) because presence of free electrons causes most metals to reflect light.

Ionic Bonding

- occurs between *metals* and *non-metals*..
  - **Metal** atoms have a *low number of valence electrons* and a *low electronegativity*.

  - **Non-metal** atoms have *numerous valence electrons*.

Ionic Bonding

- If the electron clouds overlap (bond)..

  Metals

  – lose valence electrons
    achieve a stable valence shell (usually 8 e-)
    gains a positive charge, ie a positive ion.

- **Non-metals**
  – gain valence electrons
    achieve a stable valence shell (usually 8 e-)
  – Will gain a negative charge ie. a negative ion

Formation of positive ions

- *eg:* The charge is the number of valence e- it has to lose.
  - Na atom [Ne] $3s^1$ loses one electron to become a sodium ion(Na^{+1}) with [He] $2s^2 2p^6$configuration
  - K atom [Ar] $4s^1$ loses one electron to become a potassium ion (K^{+1}) with [Ne] $3s^2 3p^6$ configuration
  - This is common for ALL Group 1 elements

Formation of positive ions

A Mg atom [Ne]$3s^2$ loses 2 electrons to become a magnesium ion (Mg^{+2}) with [He] $2s^2 2p^6$ configuration

A Ca atom [Ar] $4s^2$ loses 2 electrons to become a calcium ion (Ca^{+2}) with [Ne] $3s^2 3p^6$ configuration

This is common for ALL Group 2 elements

Formation of positive ions

- Al atom [Ne] $3s^2 3p^1$ loses 3 electrons to become an aluminum ion (Al^{+3}) with [He] $2s^2 2p^6$ configuration.
- This is common for Group 13 metals ONLY
- General note: metals will always make positive ions

Formation of negative ions

- F atom [He] $2s^2 2p^5$ goes to the fluoride ion (F^{-1}) with [He] $2s^2 2p^6$
- Cl atom [Ne]$3s^2 3p^5$ goes to the chloride ion (Cl^{-1}) with [Ne] $3s^2 3p^6$

Formation of positive ions

- O atom [He] $2s^2 2p^4$ goes to the oxide ion (O^{-2}) with [He] $2s^2 2p^6$
- The charge is the number of electrons needed to get 8 in the outermost level and make the number negative.
  - Ex: O gains 2 e^- and has –2 charge
- Nonmetals will **usually** make negative ions

Formation of ionic compounds

- *The Positive and negative ions formed by the gaining and losing electrons will*
  - attract each other to form a three dimensional continuous lattice structure (or crystal.)
  - Each positive ion is surrounded by a number of negative ions.
  - Each negative ion is surrounded by a number of positive ions.
  - The ratio of positive to negative ions in the lattice is determined by the charges of the ions

Properties of Ionic compounds and why they have these properties

- 1) Melting point and boiling point are high because a large amount of thermal energy is required to separate the ions which are bound by strong electrical forces.
- 2) Electrical conductivity is poor because there are no free electrons.

Properties of Ionic compounds and why they have these properties

- 3) The crystals of ionic compounds are hard because the ions are bound strongly to the lattice and aren't easily displaced.
- 4) The crystals of ionic compounds are brittle because distortion cause ions of like charges to come close together then sharply repel.

Covalent Bonding

- Bonding between *non-metals* and *non-metals*.
  - Therefore all atoms included have *fairly high electronegativity* and *few vacancies in valence energy levels*.
  - When they bond, they gain electrons to achieve stable configuration.
  - Hence, electrons are **shared**.

Covalent Bonding

- Sharing produces low energy (stable) electron arrangements.
- *ie:*
  - Full outer shell (eg He $1s^2$; Ne $1s^2\ 2s^2\ 2p^6$)
    8 electrons (4 pairs) in outer shell (eg Ar $1s^2\ 2s^2\ 2p^6\ 3s^2 3p^6$)

When more than one pair of electrons are shared

- There some situations in covalent bonding where there are insufficient pairs of electrons.
- When this occurs, atoms will share multiple pairs of electrons.
- carbon is one of the elements that does this quite readily.

Multiple covalent bonds

- When two atoms share 2 pairs of electrons, the bond is called a double covalent bond
- When two atoms share 3 pairs of electrons, the bond is called a triple covalent bond.

The Octet Rule

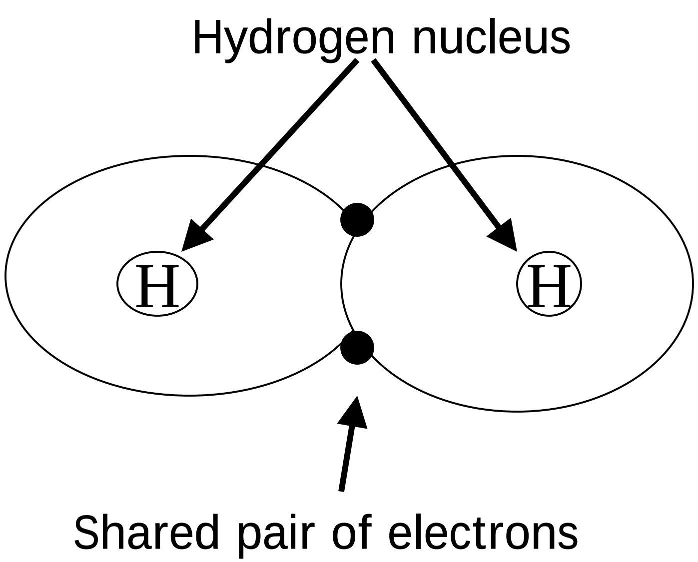
- The **octet rule** says that atoms tend to gain, lose or share electrons so as to have eight electrons in their outer electron shell.
- It is a very useful rule but you should also know that there are *many* bonding situations where it does *not* apply.
- As you learn to use the octet rule, also learn to recognize situations where it does not apply and disregard it in those situations.

Covalent bonding

- **Covalence** is the number of electrons an atom needs to produce a stable outer shell. The number of *shared pairs (covalent bonds)* of electrons an atom forms. eg Hydrogen *H* needs 1 additional electron

Covalent bonding in the hydrogen molecule

- Hydrogen bonds1 pair of e-'s shared between 2 atoms - *covalent bond*.



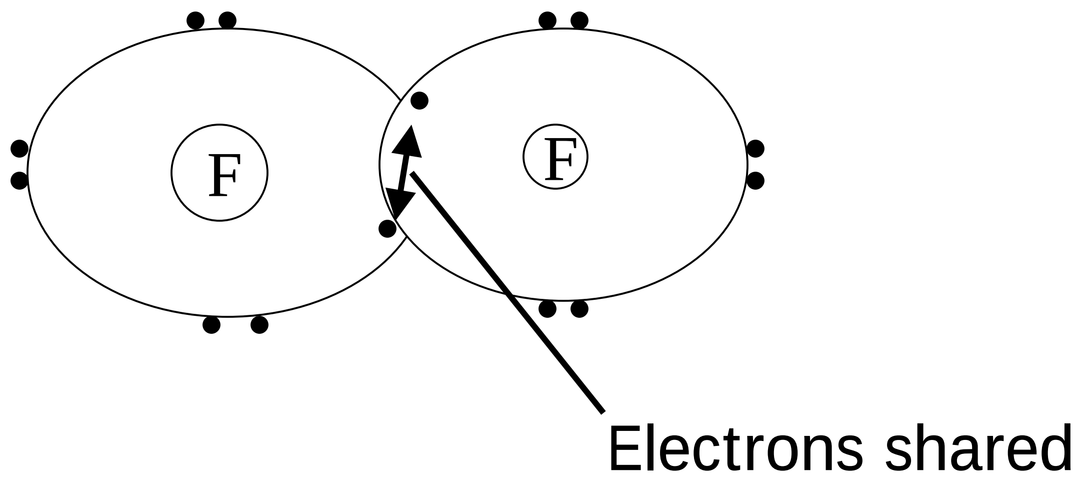
Covalent bonding in the hydrogen molecule

- Hydrogen molecule consists of 2 covalently bonded hydrogen atoms, which have no tendency to bond further (both have achieved a stable outer shell).
- Each molecule exists independently

- Bonding pairs of electrons orbits both nuclei - attracts both nuclei - provides bonding force

Covalent bonding in other diatomic molecules

- eg Fluorine F [He] $2s^2 2p^5$ needs one electron to achieve Noble gas configuration.



Properties of covalently bonded compounds

- 1) Do not conduct electricity because Electrons are tightly bound to atoms or shared by atoms in covalent bonds and do not move.
- 2) Melting and boiling points low because Forces of attraction between molecules are weak and little thermal energy is required to separate them.

Properties of covalently bonded compounds

- 3) Soft because molecules are weakly attracted to each other and are easily displaced.

Not ALL electrons are shared equally in a covalent bond

- In covalent bonds, atoms of equal electronegativity the electron pair is shared equally between the atoms
- If the atoms in a covalent bond have differing electronegativities, the atoms with the higher electronegativity has >50% of the shared pairs of electrons and the atoms with low electronegativity has <50% of the shared pairs of electronegativity.

Polar covalent bonds

- The atom tending to gain electrons acquires a slight *negative charge* (delta negative)
- The atom tending to lose electrons acquires a slight *positive charge* (delta positive)
- *The bond is* **POLAR.**

Expansion of the Octet

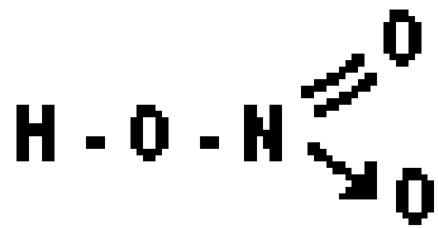
- Atoms, which have room for more than 8 electrons in their outer shell, may form covalent bonds, which result in 10, 12, or 14 outer shell electrons (even numbers are stable).

- ie atoms in Period 3 or below are able to form shared pairs -> even number of e- results
  - (room for 18 valence e-)
- eg:
  - P $[Ne]3s^2 3p^3$ - 3 shared pairs, 8 valence e-'s, or 5 shared pairs, 10 valence e-'s.

  - S $[Ne]3s^2 3p^4$ - 2 shared pairs, 8 valence e-'s, or 4 shared pairs, 10 valence e-'s, or 6 s.p, 10 v.e-'s.

Co-ordinate Bonds

- A covalent bond where the shared pair of e-'s is provided by 1 atom.
  It is represented by an arrow, showing the direction in which the shared e-'s are provided
  - eg HNO_3

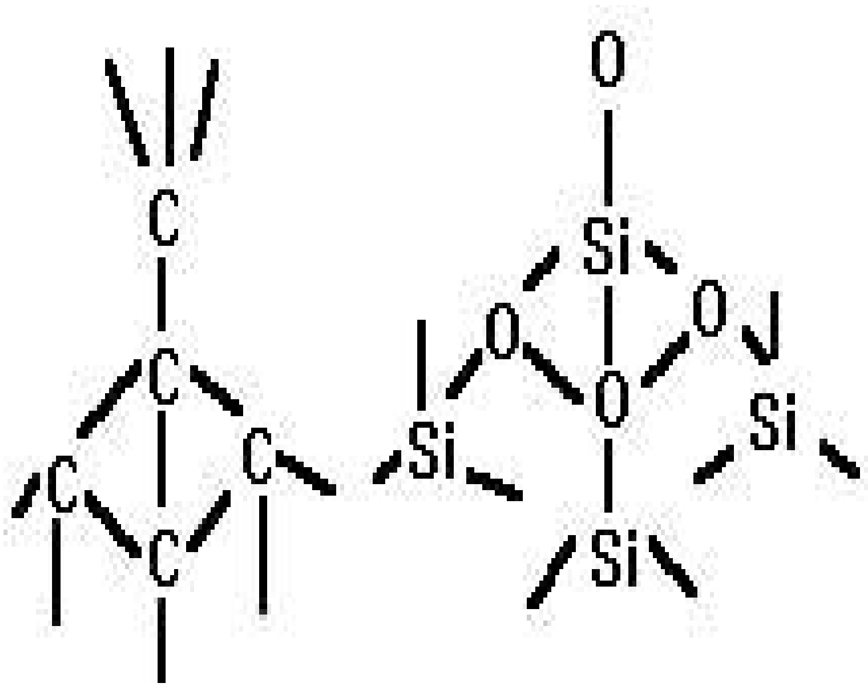


  - Co-ordinate bonds aren't very common. They usually occur in atoms that cannot expand the octet.

Covalent Network (Lattice)

- Bonding between non-metals.
  Form covalent bonds (shared electron pairs).
- Do not form separate (discrete) molecules but a *continuous network*.

Examples of covalent crystals



- eg Diamond (carbon) -> - Silicon Dioxide SiO_2.

Properties of Covalent crystals

1) Poor Electrical conductivity because electrons are held either on the atoms or within covalent bonds.
   They cannot move through the crystal.
   Graphite is an exception.
2) Very high melting points because each atom is bound by strong covalent bonds.
   Many covalent bonds must be broken if the solid is to be melted and a large amount of thermal energy is required for this.

Properties of Covalent crystals

- 3) These compounds are Hard because the atoms are strongly bound in the crystal, and are not easily displaced.
- 4) Covalent network substances are brittle.

  If sufficient force is applied to a crystal, covalent bond are broken as the lattice is distorted.

  Shattering occurs rather than bending of the shape.

Resonance

- There are a number of compounds and polyatomic ions that cannot be written using one single structure.
- Linus Pauling developed what today is called "resonance theory."

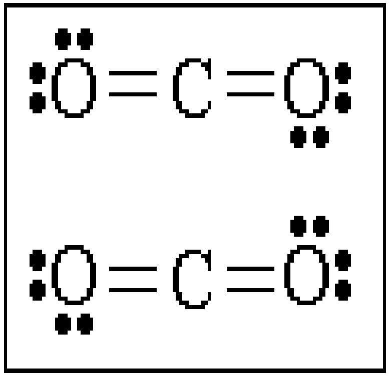
Resonance

- **Resonance happens when more than one valid Lewis dot-diagram (or what Pauling calls a valence-bond structure) can be written for a molecule or ion.**
- **When this happens, the true structure is a blend of all the different possible structures.**

Resonance

- Here is another example, using the molecule CO_2:

Carbon Dioxide Resonance Structures

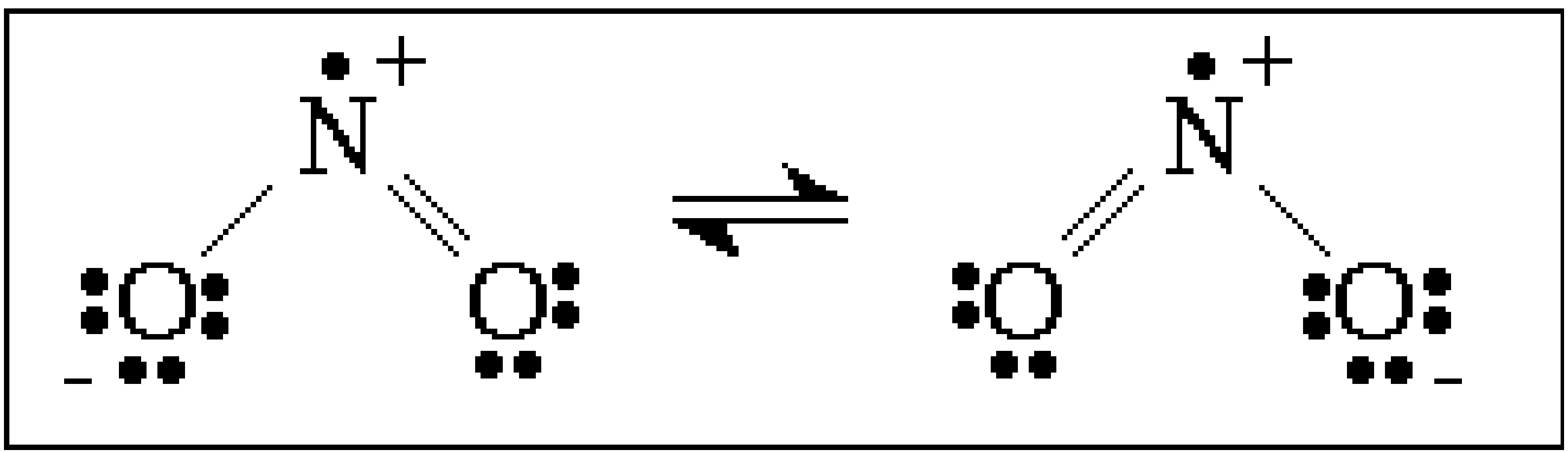


Resonance

- OK, here's the deal. Neither one of those two structures really does exist.
- The real molecule that exists in nature is a "resonance hybrid" between the two.
- The real molecule acts as if it had one and one-half bonds between each of the two structures.
- The two structures above are merely descriptive aids and, in fact, never exist.

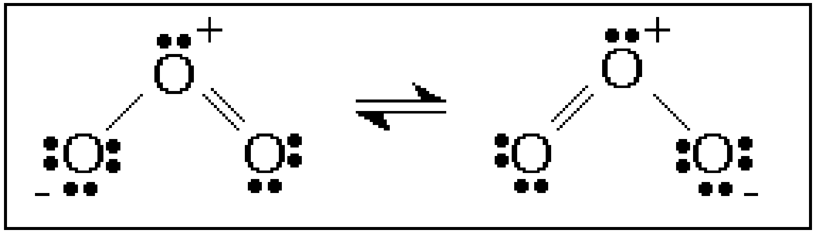
Resonance

- An example of resonance with NO_2:



Resonance

Ozone Resonance Structures



Intermolecular Forces

- Ionic Compounds and Metals
  - **Electrostatic forces -** these forces occur between charged species and are responsible for the extremely high melting and boiling points of ionic compounds and metals.

Intermolecular Forces

Covalent Compounds

London forces - all molecules have the capability to form London forces.

These are the only types of forces that non-polar covalent molecules can form. They result from the movement of the electrons in the molecule, which generates temporary positive and negative regions in the molecule.

Dipole-dipole forces

- **Dipole-dipole forces -** only polar covalent molecules have the ability to form dipole-dipole attractions between molecules. Polar covalent molecules act as little magnets; they have positive ends and negative ends that attract each other.

Hydrogen bonding

- **Hydrogen bonding -** these occur between polar covalent molecules that possess a hydrogen bonded to an extremely electronegative element, specifically - N, O, and F.
- Water is a major example of this type of bonding and explains why the mp and bp of water are so high.

Hybridization

Molecular Geometry

- The way that some atoms covalently bond together will cause them to have a particular molecular shape.
- There are three theories that attempt to explain this:
  - VESPER theory
  - Valence Bond Theory
  - Molecular orbital theory